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ALCHEM 20

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alberta chemistry project materials

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Forward

HEM (Alberta Chemistry Project Materials) was initiated by the Edmonton Regional Chemistry Council. Funding has been provided by the Edmonton Public School Board and to a lesser extent by the Edmonton Separate School Board. ALCHEM is a local curriculum materials development project aimed at preparing classroom oriented materials. The unique format has received widespread approval both by students and teachers. The materials have been pilotted by over 20 000 students and 100 teachers. The feedback has been very positive. On a comparative study, ALCHEM students have significantly outperformed students using other materials. On the same comparative study ALCHEM students showed an equal or higher interest in chemistry.

The main objectives when writing the materials have been:

1. keep a high level of chemistry
2. provide easy-to-learn from classroom materials
3. integrate applied chemistry and theoretical chemistry

The applied chemistry has largely centered around:

1. environmental chemistry
2. consumer chemistry
3. industrial chemistry
4. history of chemistry

Acknowledgements

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CHEMISTRY 20
COURSE OUTLINE

Unit F: Review of Chemistry 10

1. Elements and the periodic table
2. Bonding and nomenclature
3. Chemical reactions
4. The mole concept
5. Gravimetric stoichiometry

Unit G: Chemical Bonding

1. Chemical bonding between atoms including covalent, network covalent, ionic and metallic bonding
2. Shapes of molecules
3. Electronegativity, polarity of bonds and polarity of molecules
4. Intermolecular bonding

Unit H: Organic Chemistry

1. Bonding
2. Nomenclature
3. Molecular shapes
4. Properties
5. Common sources and uses

of

- alkanes
- alkenes
- alkynes
- aromatics
- organic halides
- alcohols
- carboxylic acids
- esters

Unit I: Solutions

1. Basic problems involving solution concentration and solution volume and dilution
2. Pipetting techniques
3. Preparation of standard solutions
4. Electrolytes and nonelectrolytes
5. Dissociation and concentration of ionic species
6. Saturated solutions and solubility

Unit J: Reactions in Solutions

1. Types of reactions in solution
2. Writing ionic equations
3. Stoichiometric calculations involving aqueous solutions
4. Titration and filtration techniques

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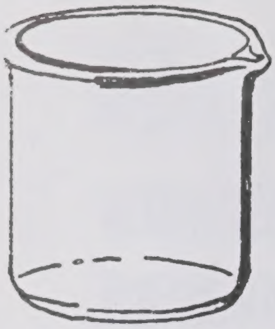
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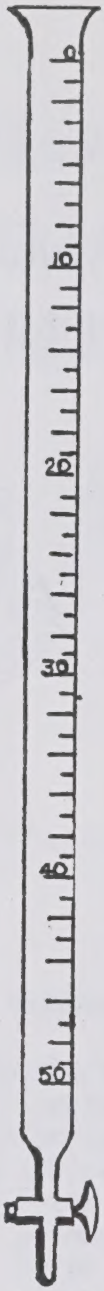
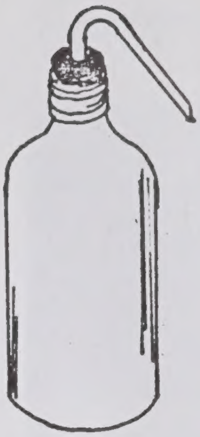


buret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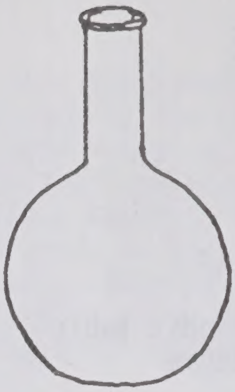
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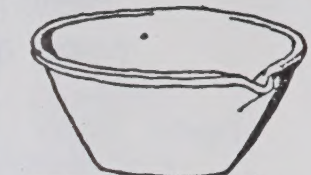
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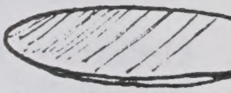
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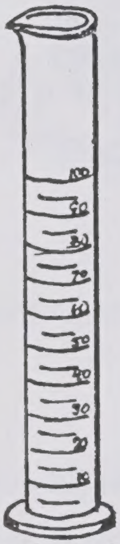
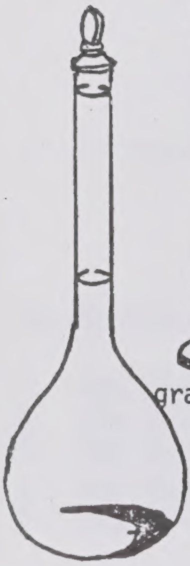


graduated cylinder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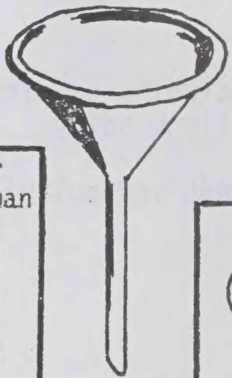
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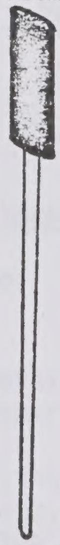
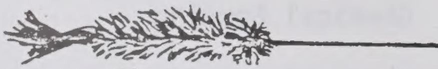

rubber policeman

stirring rod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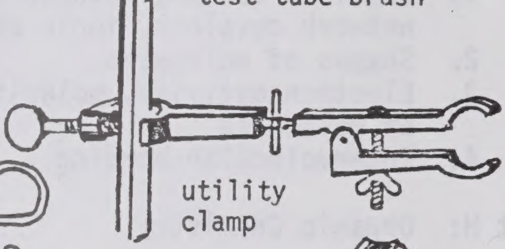
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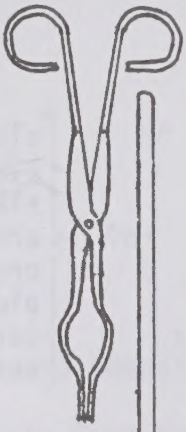
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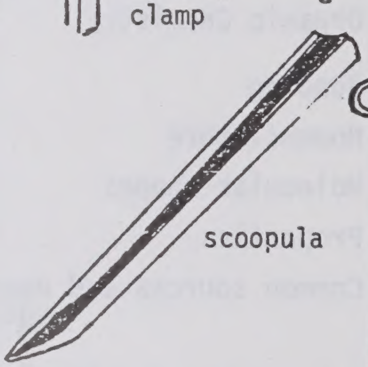
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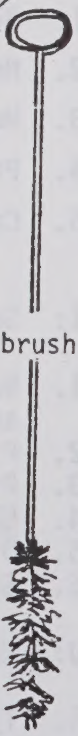
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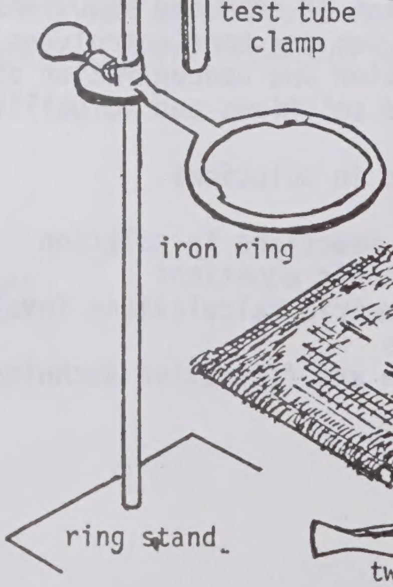
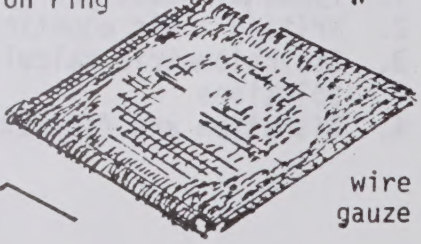

test tube clamp

iron ring

ring stand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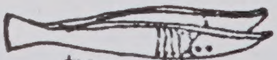
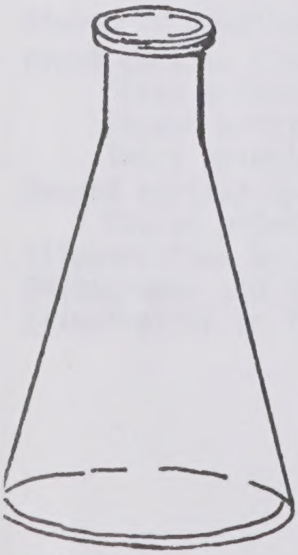
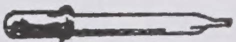

tweezers


Erlenmeyer flask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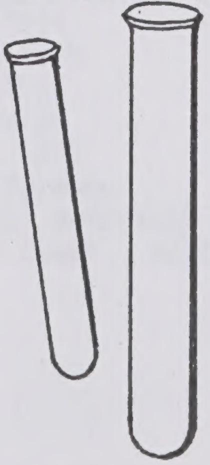
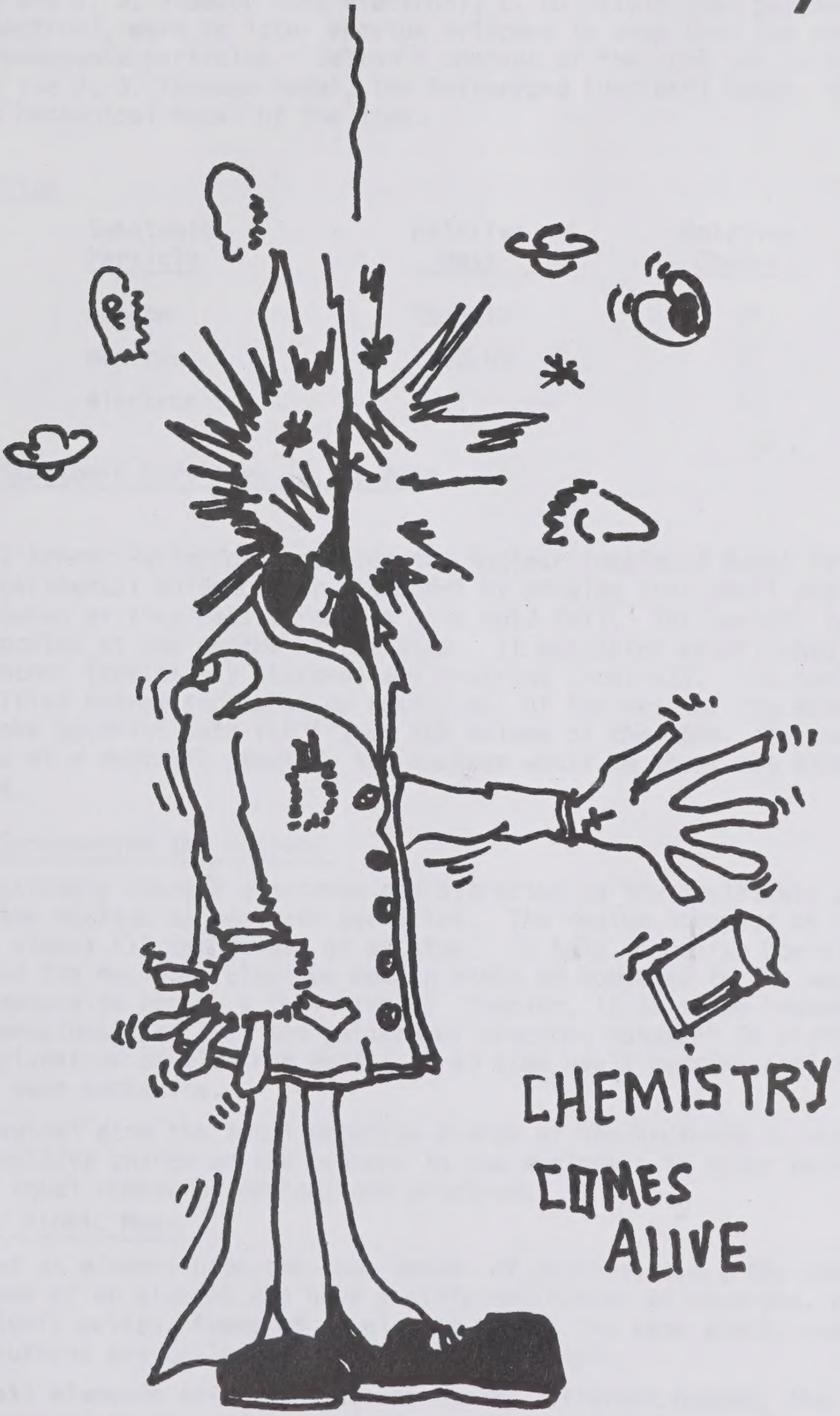

test tubes

pipet

UNIT F

Review of Chemistry 10



REVIEW OF CHEMISTRY 10
ATOMIC STRUCTURE

Models of the Atom

In 1804 John Dalton used a table of atomic masses as evidence for the theory that all substances are composed of atoms - the smallest particles of matter. Scientists such as Michael Faraday and J. J. Thomson (the electron), E. Goldstein (the proton), and James Chadwick (the neutron), were to later provide evidence to show that the atom is composed of three basic subatomic particles. Dalton's concept of the atom was to be progressively refined through the J. J. Thomson Model, the Rutherford (nuclear) Model, the Bohr Model and now the Quantum Mechanical Model of the atom.

Subatomic Particles

Subatomic Particle	Relative Mass	Relative Charge
proton	1836.12	1+
neutron	1838.65	0
electron	1	1-

Arrangement of Subatomic Particles in the Atom

1. The Nucleus

In 1912 Ernest Rutherford proposed the Nuclear (nucleus) Model for atoms. Rutherford provided experimental evidence for the model by showing that small positive particles are widely scattered as they pass through a thin gold foil. The nucleus is a small massive structure located at the centre of the atom. It was later established that the nucleus contains protons (positively charged) and neutrons (neutral). The nucleus of an atom has all the positive charge and makes up nearly all of the mass of the atom, but occupies only about one quadrillionth (10^{-15}) of the volume of the atom. By analogy, if an atom was the size of a football stadium, the nucleus would be about the size of an ant at centre field.

2. The Region Surrounding the Nucleus

The negatively charged electrons are attracted by the positively charged nucleus and move about the nucleus as separate particles. The region occupied by the moving electrons constitutes almost all the volume of an atom. To help visualize how electrons occupy the region around the nucleus, electron motion might be compared to the way a spinning propellor appears to occupy a full circle. However, it is to be remembered that an atom is three dimensional and that the pattern of electron movement is highly complex. A more complete explanation of electron motion in an atom would require applying the complex theories of wave mechanics.

In a neutral atom the total negative charge of the electrons is exactly balanced by the total positive charge of the protons in the nucleus. In other words, a neutral atom contains an equal number of protons and electrons.

3. Isotopes and Atomic Mass

Atoms of an element have the same number of protons; i.e., the same *atomic number*. However, atoms of an element can have a different number of neutrons, and consequently different atomic masses. Atoms of an element having the same atomic number but different number of neutrons are called *isotopes* of that element.

Since all elements exist as isotopes having different masses, the *atomic mass* of an element represents the *average relative mass* of all the naturally occurring isotopes of that element. All atomic masses are relative to a carbon isotope with six protons and six neutrons.

4. Changes in Atoms During Chemical Reactions

In chemical reactions a rearrangement of only the electrons takes place - the nucleus is not affected. Understanding chemistry depends upon understanding how the electrons of atoms are arranged around their nuclei and upon how the electrons of atoms interact with the electrons of other atoms.

When atoms approach each other closely their electrons become simultaneously attracted by the positive nuclei of other atoms. The simultaneous attraction for the same electrons by the nuclei of two or more atoms causes electron rearrangements among atoms. The Electron rearrangements may be considered to be of two distinct types:

(1) *the loss and gain of electrons*, and (2) *the sharing of electrons*

Arrangement of Electrons Around the Nucleus - The Bohr Model

In 1912-13 Neils Bohr's study of the atomic spectra of hydrogen showed that the electrons in atoms move around the nucleus in specific energy states (or energy levels). Bohr proposed that electrons of specific energy moved in definite orbits around the nucleus and that the electrons could not exist between orbits. Bohr's theory of electron motion may be compared to the motion of planets around the sun, or to the layered structure of an onion. It should be realized, however, that electron motion within an atom is much more complex than either of the preceding analogies would indicate.

The concept of electrons located in specific energy levels is still retained in the present theory of the atom. The main difference between the Bohr and the newer Quantum Mechanics Theory is that electrons are not considered to be revolving around the nucleus in orbits, but to occupy "orbitals". The orbitals represent probable patterns of movement of electrons around the nucleus which are characteristic of the energies of the electrons. Therefore, according to the newer ideas, an electron is described in terms of its orbital rather than in terms of its orbit.

The electrons of highest energy are the ones that are (usually) involved in chemical reactions. These *outer energy level (outermost)* electrons are called *valence (bonding) electrons*.

The Periodic Table

The periodic table is a systematic way of organizing the elements. In this table the elements are arranged horizontally in order of increasing atomic numbers to give seven horizontal *periods*. Each period a) starts with an alkali metal and ends with a noble gas (with the exception of the first period), and b) corresponds to the energy level of the valence electrons in the element represented.

The arrangements of atoms in horizontal rows in order of increasing atomic number forms vertical columns of elements having identical or similar valence-electron-structures and thus similar chemical properties. These vertical columns are known as *groups or families* of elements.

The heavy staircase line starting at boron and staggered down the table separates the elements into metals and nonmetals. The elements to the right of the line are nonmetallic whereas those to the left (except H and Be) are metallic. With some exceptions, the electronic structure of metals is such that their atoms have 1, 2 or 3 valence electrons while the nonmetals have 5, 6 or 7 valence electrons. Thus the arrangements of elements in the periodic table is such that nonmetallic character increases within a period with increasing atomic number (i.e., from left to right), and metallic character increases within a group with increasing atomic number (i.e., from top to bottom).

REVIEW OF CHEMISTRY 10
ATOMIC STRUCTURE

Simple Ions

In chemical reactions Group A atoms may lose or gain electrons to acquire the more stable electron structure of the nearest noble gas - usually this means a structure with eight electrons in the outer energy level (i.e., an octet structure) or a structure of two electrons in the first energy level. There is no simple rule that can be used to predict how all ions are formed, however, it can be generalized that many common ions are formed in the following way:

> *An atom that has an electron population close to that of an inert gas may gain or lose electrons so as to acquire the same number of electrons as that inert gas.*

The loss or gain of electrons gives the atom a more stable electronic structure but unbalances the number of positive or negative electrical charges in such a way that the atom becomes either positively or negatively charged. Atoms that acquire electrical charge in this manner are called *ions*.

Electron Energy Level Representation for Atoms and Ions

As pointed out by the Bohr Theory, electrons of atoms are located in certain specific energy levels about the nucleus. The successive energy levels are also referred to as "electron shells" and may contain only a limited number of electrons. The maximum number of electrons in any specific energy level is related to the general arrangement of atoms in the periodic table. The period number corresponds to the outermost energy level in which electrons are located in the neutral atom. For example, for the first three periods the period one elements contain electrons in the first energy level only; period two elements contain electrons in the first and second levels; period three elements contain electrons in the first, second and third levels. The sequence of locating the last electrons in the outermost energy level applies not only to the elements of the first three periods but also to all Group A elements in all periods.

Ionic charges for the Group A elements can be readily determined from their group number. The Group A elements with fewer than four electrons in their outermost energy level (i.e., with group number less than four) ten to lose electrons to acquire the more stable structure of the nearest inert gas. The atoms losing electrons (losing negatively charged particles) acquire a positive charge numerically equal to the number of electrons lost. On the other hand, the Group A elements with more than four electrons in their outermost energy level (i.e., with group number greater than four) tend to gain electrons to acquire the more stable structure of the nearest inert gas. The atoms gaining electrons (gaining negatively charged particles) acquire a negative charge numerically equal to the number of electrons gained.

Since the last electron added in the Group B elements (the transition elements) is located in an inner energy level, their ionic charges cannot be reliably obtained from their group number. For the transition elements the positive charge given in the periodic table will be assumed.

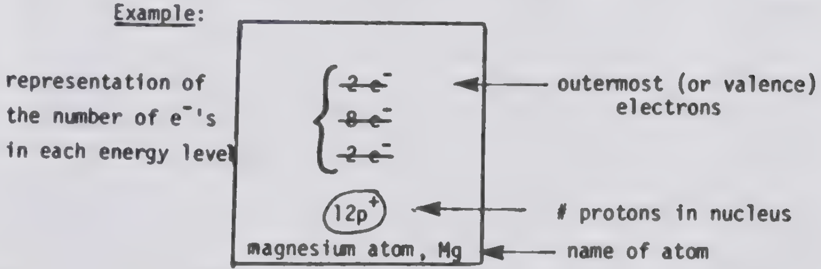
In summary, atoms form ions by losing or gaining electrons, which results in an imbalance in the number of protons and electrons. Since the number of protons in an atom remains fixed during chemical reactions the loss or gain or electrons determines ionic charge.

REVIEW OF CHEMISTRY 10

ELECTRON ENERGY LEVEL REPRESENTATION FOR ATOMS

Complete the following abbreviated periodic table for atoms as shown in the example.

Example:


representation of the number of e⁻'s in each energy level
2 e⁻
8 e⁻
2 e⁻
12p⁺
magnesium atom, Mg
outermost (or valence) electrons
# protons in nucleus
name of atom

Group IA	IIA	IIIA	IVA	VA	VIA	VIIA	VIIIA

QUESTIONS:

1. What is the relationship between group number and number of outermost (or valence) electrons?

____

2. What is the relationship between period number and the number of energy levels in which electrons are accommodated?

____

3. According to the above abbreviated periodic table, how many electrons can be accommodated before a new energy level is started in each of the first three energy levels?

First energy level ____________ Second energy level ________________ Third energy level____________

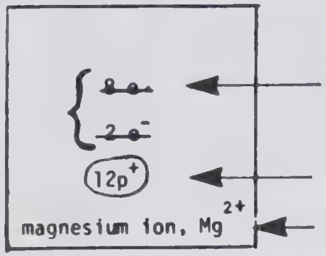
REVIEW OF CHEMISTRY 10

ELECTRON ENERGY LEVEL REPRESENTATION FOR IONS

Complete the following abbreviated periodic table for ions as shown in the example.
For Group VIIIA (noble gases) give the electron energy level representation for the neutral atoms.

__Example__:

representation of the number of e^-'s in each energy level



energy level structure characteristic of nearest noble gas

protons in nucleus

name of ion

Group IA	IIA	IIIA	IVA	VA	VIA	VIIA	VIIIA

__QUESTIONS__:

1. What relationship exists between the electron structure of a group A ion and the electron structure of the nearest noble gas?

   ____

2. What problem arises when trying to predict the charge on ions in group IVA?

   ____

3. What charge do the ions from the following groups assume?

   Groups IA ________ Group IIA ________ Group IIIA ________ Group VA ________ Group VIA ________ Group VIIA ____________

REVIEW OF CHEMISTRY 10
PERIODIC TABLE

MATCHING EXERCISE

(Use the periodic table for the following exercise)

1. Vertical columns in the periodic table ________
2. Horizontal rows in the periodic table ________
3. Number of protons in the nucleus of an atom ________
4. Group number of the alkali metals ________
5. Group number of the inert (noble) gases ________
6. Group number of the halogens ________
7. Group number of elements with four valence electrons ________
8. Pure substance containing atoms of only the same atomic number ________
9. The most chemically reactive metal ________
10. The most chemically reactive nonmetal ________

a. atomic number
b. elements
c. fluorine
d. francium
e. groups (or families)
f. group IA
g. group IVA
h. group VIIA
i. group VIIIA (or 0)
j. periods

Describe the following gradations of properties within the periodic table:

1. (a) The gradation of metallic character with increasing atomic number in a period.

____

____

(b) The gradation of metallic character with increasing atomic number in a group.

____

____

2. The attraction by the nuclei of atoms for their valence electrons with increasing atomic number in a period.

____

____

REVIEW OF CHEMISTRY 10
ATOMIC STRUCTURE

Complete the following table:

	SPECIES	SYMBOL	# PROTONS	# ELECTRONS	NET CHARGE
	EXAMPLE: sodium ion	Na^+	11	10	1+
1.	oxygen atom				
2.	fluoride ion				
3.		C			
4.		Cl^-			
5.			12		2+
6.			16		2-
7.				18	1+
8.				10	0
9.		Ba^{2+}			
10.	helium atom				
11.		H^+			
12.			7	10	
13.		Fe^{3+}			
14.		Sn^{4+}			
15.	calcium ion				
16.		Al^{3+}			
17.			29		2+
18.				54	1-
19.	gold atom				
20.				54	1+

REVIEW OF CHEMISTRY 10

SUMMARY - DETERMINING CHARGES ON IONS

Simple Metallic Ions (Simple Cations)

Group Number of Metal	Charge on Metallic Ion
Group IA	1+
Group IIA	2+
Group IIIA	3+
Group IVA	4+

The charges assumed by transition metals and by metals in groups VA, VIA and VIIA can be read directly from the Periodic Table, where the charges are designated as "oxidation states". The most stable oxidation state (charge on ion) is given in bold type.

Complex Metallic Ion (Complex Cation)

Only one complex metallic ion will be considered, namely the ammonium ion, NH_4^+.

Simple Nonmetallic Ions (Simple Anions)

Group Number of Nonmetal	Charge on Nonmetallic Ion
Group VIIA	1-
Group VIA	2-
Group VA	3-

Note: The periodic table lists several oxidation states for many of the nonmetals. For purposes of this course, consider that nonmetallic ions tend to only assume the negative charges given in the table above. The charges given in the table are consistent with the nonmetals acquiring the electron population of the nearest inert gas.

Complex Nonmetallic Ions (Complex Anions)

The charges on nonmetallic complex ions are given in the table "Complex Ions". Some of the nonmetallic ions are in common usage and should be memorized. The formula and name of the following acids and their associated complex ions should be memorized:

$H_2SO_{4(aq)}$	sulfuric acid
$HNO_{3(aq)}$	nitric acid
$HClO_{3(aq)}$	chloric acid
$H_3PO_{4(aq)}$	phosphoric acid
$H_2CO_{3(aq)}$	carbonic acid
$CH_3COOH_{(aq)}$	acetic acid

The formula and name of less common complex anions can be obtained from the table provided.

REVIEW OF CHEMISTRY 10
NAMING AND WRITING FORMULAE FOR IONIC COMPOUNDS, ACIDS AND MOLECULAR SUBSTANCES

Complete the following tables.

Ionic Compounds

	Formula	Name of Compound
1.	Na_2O	
2.		strontium sulfite
3.		gallium sulfide
4.	NH_4CH_3COO	
5.		zinc nitrate
6.	$NiBr_3$	
7.		magnesium acetate
8.	$Ni_2(SO_4)_3$	
9.		stannic oxide
10.	$CaBr_2$	

	Formula	Name of Compound
11.		ammonium carbonate
12.	$PbSO_4$	
13.		iron(II) sulfate
14.		zinc chlorate
15.	ZnO	
16.	$KBrO_3$	
17.		cupric oxide
18.	$NaHCO_3$	
19.	PbO_2	
20.		chromium(III) chloride

Acids

	Formula	Name of Acid
1.	$HCl(aq)$	
2.		bromic acid
3.	$HBr_{(aq)}$	
4.		hydrosulfuric acid
5.	$H_2SO_{4(aq)}$	
6.		hydriodic acid

	Formula	Name of Acid
7.	$HNO_3(aq)$	
8.		phosphoric acid
9.	$HClO_{3(aq)}$	
10.		carbonic acid
11.	$CH_3COOH_{(aq)}$	
12.		iodic acid

Molecular Substances

	Formula	Name of Compound
1.	SiH_4	
2.		bromine gas
3.	AsF_3	
4.		carbon tetrachloride
5.		hydrogen sulfide
6.	SO_2	
7.		iodine
8.	NCl_3	
9.		dinitrogen pentoxide
10.	NH_3	

	Formula	Name of Compound
11.	N_2	
12.		bromine monochloride
13.	PH_3	
14.		carbon disulfide
15.		sulfur trioxide
16.	CH_4	
17.		dinitrogen trioxide
18.	P_4O_{10}	
19.		sucrose
20.	SF_6	

REVIEW OF CHEMISTRY 10
CHEMICAL NOMENCLATURE

Complete the following table:

	Chemical Formula	Name of Compound
1.	$Al(OH)_3$	
2.		sodium sulfate
3.		sodium nitrate
4.	$Al_2(SO_4)_3$	
5.		calcium chloride
6.	NH_4OH	
7.		hydrogen peroxide
8.	N_2O	
9.		methane
10.	$H_2SO_{4(aq)}$	
11.	H_3PO_4 (aq)	
12.		boric acid
13	$(NH_4)_2SO_4$	
14.	SnF_2	
15.		carbonic acid
16.	PbO_2	
17.		silicon dioxide
18.	$NaClO$	
19.		potassium permanganate
20.	KNO_3	
21.	K_2CO_3	
22.		hydrofluoric acid
23.	C	
24.		sodium hydroxide
25.	$NaHSO_4$	

	Chemical Formula	Name of Compound
26.		magnesium sulfate
27.	$Ca(OH)_2$	
28.		sodium thiosulfate
29.	CaO	
30.		copper(II) sulfate
31.		sulfur
32.	B_2H_6	
33.	KI	
34.		phosphorus
35.	SO_3	
36.		sodium chlorate
37.	Na_2SiO_3	
38.		methanol
39.		chloric acid
40.		plumbous sulfate
41.	$Ca(HCO_3)_2$	
42.		nitrogen trichloride
43.		sodium bisulfite
44.	CS_2	
45.	$H_2S_{(aq)}$	
46.		silicon carbide
47.		aluminum phosphate
48.		hydrogen nitrate
49.	SiC	
50.	N_2H_4	

Note:

It is suggested that the balance of the review materials be left until after the completion of Unit G, Chemical Bonding and Unit H, Organic Chemistry. The remaining review materials have more direct application to Unit I, Solutions and Unit J, Reactions in Solution.

REVIEW OF CHEMISTRY 10
BALANCING EQUATIONS AND IDENTIFYING REACTION TYPE

Balance the following equations by writing in the simplest whole number coefficients in the space provided. Identify the reaction type in the space to the right of the equation.

Reaction Type

1. ____ $A\ell$ + ____ O_2 $\rightarrow$ ____ $A\ell_2O_3$ ____________

2. ____ $HC\ell_{(aq)}$ + ____ $Ca(OH)_2$ $\rightarrow$ ____ HOH + ____ $CaC\ell_2$ ____________

3. ____ CH_4 + ____ O_2 $\rightarrow$ ____ CO_2 + ____ H_2O ____________

4. ____ Zn + ____ $Pb(CH_3COO)_2$ $\rightarrow$ ____ Pb + ____ $Zn(CH_3COO)_2$ ____________

5. ____ SO_3 + ____ H_2O $\rightarrow$ ____ H_2SO_4 ____________

6. ____ HgO $\rightarrow$ ____ Hg + ____ O_2 ____________

7. ____ $CaCO_3$ $\rightarrow$ ____ CaO + ____ CO_2 ____________

8. ____ NaI + ____ $Pb(NO_3)_2$ $\rightarrow$ ____ PbI_2 + ____ $NaNO_3$ ____________

9. ____ $C\ell_2$ + ____ NaI $\rightarrow$ ____ I_2 + ____ $NaC\ell$ ____________

10. ____ $A\ell_2(SO_4)_3$ + ____ $Ca(OH)_2$ $\rightarrow$ ____ $A\ell(OH)_3$ + ____ $CaSO_4$ ____________

11. ____ $A\ell_2(SO_4)_3$ + ____ $Ca(HCO_3)_2$ $\rightarrow$ ____ $A\ell(OH)_3$ + ____ $CaSO_4$ + ____ CO_2 ____________

12. ____ C_8H_{18} + ____ O_2 $\rightarrow$ ____ CO_2 + ____ H_2O ____________

13. ____ $NaHCO_3$ + ____ H_2SO_4 $\rightarrow$ ____ Na_2SO_4 + ____ H_2CO_3 ____________

14. ____ CO_2 + ____ H_2O $\rightarrow$ $H_2CO_{3(aq)}$ ____________

15. ____ Ca + ____ HOH $\rightarrow$ ____ H_2 + ____ $Ca(OH)_2$ ____________

16. ____ $H_2SO_{4(aq)}$ + ____ $A\ell(OH)_3$ $\rightarrow$ ____ HOH + ____ $A\ell_2(SO_4)_3$ ____________

17. ____ $Ca(HCO_3)_2$ + ____ $Ca(OH)_2$ $\rightarrow$ ____ $CaCO_3$ + ____ H_2O ____________

18. ____ Na_2CO_3 + ____ $Ca(OH)_2$ $\rightarrow$ ____ NaOH + ____ $CaCO_3$ ____________

REVIEW OF CHEMISTRY 10

WRITING AND BALANCING EQUATIONS AND IDENTIFYING REACTION TYPE

From the names of the reactants given you are to:

1. indicate the reaction type
2. predict the products and write a balanced equation
3. complete the word equation

1. Reaction type: ______________________________

   Balanced equation:

   Word equation: ammonia → ________________ + ______________________

2. Reaction type: ______________________________

   Balanced equation:

   Word equation: zinc + hydrochloric acid → _________ + ________________

3. Reaction type: ______________________________

   Balanced equation:

   Word equation: aluminum + oxygen → _________________________________

4. Reaction type: ______________________________

   Balanced equation:

   Word equation: acetic acid + barium hydroxide → ________ + _____________

5. Reaction type: ______________________________

   Balanced equation:

   Word equation: iron (III) nitrate + sodium hydroxide → ________ + ___________

6. Reaction type: ______________________________

   Balanced equation:

   Word equation: mercuric oxide → _____________ + _____________________

7. Reaction type: ______________________________

   Balanced equation:

   Word equation: butane (C_4H_{10}) + oxygen → _____________ + ______________

8. Reaction type: ______________________________

   Balanced equation:

   Word equation: calcium hydroxide + ammonium chloride → _________ + ____________

REVIEW OF CHEMISTRY 10

SIGNIFICANT DIGITS

There is some degree of uncertainty in every measurement. When recording data it is important that the degree of uncertainty be shown. One method of indicating uncertainty is by the number of significant digits recorded.

DEFINITION: 1. significant digits are all of the certain digits from a measurement plus one uncertain (estimated) digit.

COUNTING: 1. count all digits from 1 to 9 plus zeros in between and following other digits.

2. do NOT count zeros in front of a value because they only serve to set the decimal place.

EXACT NUMBERS: 1. numbers that are defined. For example: 5,280 feet = 1 mile exactly

2. numbers that result from counting objects. For example: 32 students, 158 beakers, $4.95 (exactly 495 cents).

Exact numbers are not uncertain and are said to have an infinite number of significant digits.

Measured Value	No. of Sig. Digits	Measured Value	No. of Sig. Digits
15⑥ pounds	3	120.5⓪ ℓ	5
0.260② m	4	0.0501⓪ kg	4
6.0② x 10^{23} molecules	3	7.② g	2
The digits circled above are uncertain (estimated)			

ROUNDING: 1. when the first digit after those retained is less than 5, all digits retained remain the same.

2. when the first digit after those retained is 5 or greater, the last digit retained is increased by one.

Examples:

- to 2 significant digits, 12.6 becomes 13
- to 3 significant digits, 92.600,000 becomes 9.26 x 10^7 (scientific notation necessary)
- to 4 significant digits, 32.0149 becomes 32.01 (consider only the digit following the fourth when rounding; disregard the 9)
- to 3 significant digits, 0.04305 becomes 0.0431 (disregard zeros in front of the 4)

REVIEW OF CHEMISTRY 10
CALCULATIONS INVOLVING MEASURED VALUES

Since all measurements involve some uncertainty, the answers which are calculated using measured values must show the uncertainty. This uncertainty can be shown by writing the calculated answer to the correct number of significant digits. Two general rules that, in all instances, apply to calculations involving measured values are:

1. a "known" digit combined with an uncertain digit (by addition, subtraction, multiplication or division) ⟶ an uncertain digit.
2. any measured value or any calculation using measured values when expressed to the correct number of significant digits includes only one uncertain digit.

The above general rules as applied to some specific mathematical operations follow:

Addition and Subtraction

Rule: Round-off to the least number of decimal places, then add or subtract.

Example: To add: 12.26② kg (the circled digits are uncertain)
+744.④ kg

Remember: first round-off then calculate
12.3 kg
+744.④ kg
756.⑦ kg (expressed to the correct number of significant digits)

Multiplication and Division

Rule: Round-off the answer to the same number of significant digits as the value with the least number of significant digits used in the calculation.

Example:
13.5② mol (uncertain digits are circled)
2.① g/mol
①③⑤②
2 7 0 ④
2⑧③⑨② g

Answer expressed to the correct number of significant digits is 28 g.

Additional examples:

(a) 0.025 mole x 12.01 $\frac{\text{g}}{\text{mole}}$ = 0.30 g (expressed to the correct number of significant digits)

(b) 7.56 m ÷ 1000 $\frac{\text{m}}{\text{km}}$ = 0.00756 km (note that 1000 m per km is a defined and exact quantity, i.e. it can be expressed to any number of significant digits)

REVIEW OF CHEMISTRY 10

INTERNATIONAL SYSTEM OF UNITS (SI)

1. All your laboratory measurements will be in terms of the metric system - a decimal system that is used in all countries in scientific work.

2. Some SI base units are listed below:

Quantity	Unit	Symbol (no periods please)
length	metre	m
mass	kilogram	kg
volume	litre	ℓ
amount of substance	mole	mol
time	second	s
temperature	degrees Celsius	°C
	kelvins	K

3. Some of the most important prefixes used with SI units are listed below:

Prefix	Factor	Symbol	Examples
milli	$\frac{1}{1000}$ of	m	1 millilitre (ml) = $\frac{1}{1000}$ litre
centi	$\frac{1}{100}$ of	c	1 centimetre (cm) = $\frac{1}{100}$ metre
kilo	1000 times	k	1 kilogram (kg) = 1000 grams

Notes:

1. The cc is not an acceptable SI symbol. One cubic centimetre (cm^3) is equivalent to one millilitre (1 cm^3 = 1 ml).
2. A litre is slightly smaller than an Imperial (Canadian) quart.
3. A kilo is slightly larger than two pounds.
4. A kilometre is about five-eights of a mile (60 mi/hr is approximately 100 km/hr).

Convert the following SI Units - maintain the same number of significant digits in each case.

Example: 3.00 cm = 0.300 m

1. 490 cm = ____________ m
2. 58.8 mg = ____________ g
3. 4.67 cm^3 = ____________ ml
4. 3300 g = ____________ kg
5. 52.4 ml = ____________ ℓ
6. 0.054 kg = ____________ g
7. 42.3 cm^3 = ____________ ℓ
8. 7.423 ℓ = ____________ ml
9. 1423 mg = ____________ kg
10. 15.25 m = ____________ cm

REVIEW OF CHEMISTRY 10

MOLE CALCULATIONS

Complete the following table. Show your work.

Name	Formula	Molar Mass	# moles	mass
1. sodium dichromate				13.1 g
2. lead (II) acetate trihydrate	$Pb(CH_3COO)_2 \cdot 3H_2O$		0.210 mol	
3. sulfur dioxide			1.60 mol	
4.	K_2CO_3			82.8 g
5. aluminum oxide				25.5 g
6.	$Cu(ClO_3)_2$		2.50 mol	

REVIEW OF CHEMISTRY 10

MOLE CALCULATIONS

Complete the following table. Show your work as in the example.

Name	Formula	Molar Mass	# moles	mass
Example: carbon tetrachloride	CCl_4	1 C = 1 x 12.0 = 12.0 4 Cℓ = 4 x 35.5 = 142.0 154 g/mol	1.20 mol (given)	mass CCl_4 = # moles x molar mass = 1.20 mol x 154 g/mol = 185 g
1. sodium chloride			0.250 mol	
2.	KOH			4.61 g
3.	$(NH_4)_2SO_4$		0.400 mol	
4. silicon dioxide			2.25 mol	
5.	$Ca(NO_3)_2$			8.20 g

REVIEW OF CHEMISTRY 10

THE MOLE CONCEPT

Definition of the Mole

The mole is 6.02×10^{23} (Avagadro's number) of particles.

Definition of Molar Mass

Molar mass is the mass of one mole of things.

Calculation of Molar Mass

The molar mass of atoms, sometimes called atomic weight, can be read directly from the periodic table. (For this course you should round off all values from the periodic table to three digits.)

Example: Determine the molar mass of $Cu(ClO_3)_2$

$Cu(ClO_3)_2$: 1 Cu = 1 x 63.5 = 63.5 (round off 63.54 to 63.5)
2 Cl = 2 x 35.5 = 71.0 (round off 35.453 to 35.5)
6 O = 6 x 16.0 = 96.0 (round off 15.999 to 16.0)
230.5 or 231 g/mol (round off to 3 digits always, and add the proper units)

Calculation of # moles from Mass

Example: Determine the # moles of magnesium oxide in 8.06 g of the compound.

MgO: 1 Mg = 1 x 24.3 = 24.3
1 O = 1 x 16.0 = 16.0
40.3 g/mole

$$\boxed{\#\text{moles MgO} = \frac{\text{mass}}{\text{molar mass}}} = \frac{8.06\text{ g}}{40.3\text{ g/mol}} = 0.200\text{ mol}$$

Calculation of Mass from # moles

Example: Determine the mass of 0.250 mol of sodium phosphate.

Na_3PO_4: 3 Na = 3 x 23.0 = 69.0
1 P = 1 x 31.0 = 31.0
4 O = 4 x 16.0 = 64.0
164.0 or 164 g/mol

$$\boxed{\text{mass} = \#\text{ moles} \times \text{molar mass}}$$

Na_3PO_4 = 0.250 mol x 164 g/mol
= 41.0 g

REVIEW OF CHEMISTRY 10
MOLE CALCULATIONS

The City of Edmonton applies treatment methods to the raw water taken from the North Saskatchewan River that includes the addition of measured quantities of various chemicals. Complete the following exercise involving chemicals that are used in the treatment of water in Edmonton. Show your work.

	Name	Formula	Molar Mass	Quantity per Litre	
				Mass (g)	Moles
1		CaO		0.100	
2	sodium carbonate (soda ash or washing soda)				0.0000629
3	aluminum sulfate (alum)			0.00750	
4	carbon (activated charcoal)	C			0.00125

REVIEW OF CHEMISTRY 10
MOLE CALCULATIONS

	Name	Formula	Molar Mass	Quantity per Litre	
				Mass (g)	Moles
5	ammonium sulfate			0.00300	
6		$C\ell_2$			0.00000704
7	sodium chlorite	$NaOC\ell_2$		0.000500	
8	hydrofluosilicic acid	H_2Sif_6		0.000800	

Data for the above table obtained from a leaflet entitled "Our Famous Recipe for Your Favorite Drink", courtesy Edmonton Water.

REVIEW OF CHEMISTRY 10
STOICHIOMETRIC CALCULATIONS

Calculations based on quantities of substances involved in chemical reactions are known as stoichiometric calculations. Suppose two substances A and B combine to form a third substance C according to the reaction:

$$2A + B \longrightarrow 2C$$

We wish to know the mass of B (desired substance) which will react with a measured quantity of A (starting substance). The solution to this type of problem consists of four steps.

STEP 1: Write the balanced chemical equation.

$$2A + B \longrightarrow 2C$$

STEP 2: Convert mass of substance A (starting substance) to moles.

$$\text{moles of A} = \frac{\text{mass of A}}{\text{molar mass of A}} = \frac{\text{grams}}{\text{grams/mole}}$$

STEP 3: Use the coefficients in the balanced equation to determine the number of moles of B (desired substance) needed to combine with the moles of A (starting substance) calculated in Step 2.

$$\text{\# mole B} = \text{moles of A (starting substance)} \times \frac{\text{coefficient of B}}{\text{coefficient of A}}$$

or in this case,

$$\text{moles of B (desired substance)} = \text{moles of A} \times \tfrac{1}{2}$$

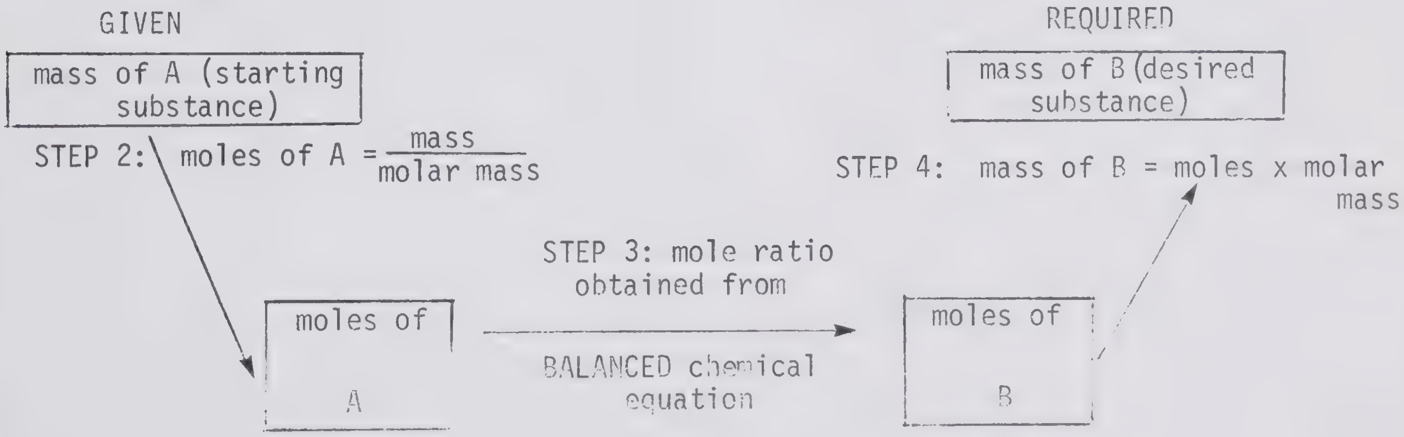
STEP 4: Convert moles of B to mass of B.

$$\text{mass of B} = \text{moles of B} \times \text{molar mass of B}$$
$$= \text{moles} \times \text{grams/mole}$$

The four general steps are summarized in the following STOICHIOMETRY FLOW CHART:

STEP 1: Write the balanced chemical equation

Gravimetric Stoichiometry Flow Diagram


GIVEN
mass of A (starting substance)
STEP 2: moles of A = mass / molar mass
REQUIRED
mass of B (desired substance)
STEP 4: mass of B = moles x molar mass
STEP 3: mole ratio obtained from
BALANCED chemical equation
moles of A
moles of B

REVIEW OF CHEMISTRY 10
GRAVIMETRIC STOICHIOMETRY

Answer the following problems - show all calculations in the space provided.

1. How many moles of each product would be formed if 0.500 mol of aluminum sulfate were to react with potassium hydroxide?

2. What mass of CO_2 gas would be produced if 4.00 mol of C_4H_8 were to react with O_2?

3. If 77.5 g of phosphorus combine with iodine, how many moles of phosphorus triodide will be formed?

4. How many grams of water will be produced when 8.42 g of propene, C_3H_6, burns in air?

5. When 0.0800 mol of nitrogen trichloride decomposes, how many moles of each product would be formed?

6. If 2.00 mol of barium chloride reacts with potassium chromate, what is the mass of each product formed?

7a. Determine the number of moles of oxygen required when 3.00 mol of propane, C_3H_8 , burns in the air.

b. Determine the number of moles of each product formed from this reaction.

8. A strip of copper is placed in an aqueous solution of silver nitrate. Determine the mass of silver that would be replaced by 0.500 mol of copper.

9. If 12.7 g of copper reacts with silver sulfate, how many moles of each product will be formed?

10. Water is decomposed by electrolysis. How many moles of hydrogen gas and oxygen gas would be formed by decomposing 4.50 g of water?

11. When aqueous solutions of lead(II) nitrate and sodium iodide are mixed, a precipitate of lead(II) iodide is formed. An excess amount of lead(II) nitrate is added to a solution containing 6.00 g of sodium iodide. What mass of lead iodide is formed?

12. What mass of silver nitrate, in solution, is needed to produce 14.0 g of silver chloride when reacted with potassium chloride?

REVIEW OF CHEMISTRY 10
LAB REGULATIONS

Laboratories contain hundreds of potentially dangerous chemicals and situations. Most lab regulations are designed to insure your safety; the rest are to ensure that your lab work will be as efficient as possible and will give accurate results. These regulations are to be studied and observed! Ignoring a lab regulation, whether deliberately or through ignorance, is cause for exclusion from the lab.

1. Study the procedure and/or teacher's preparatory notes and remarks before starting any lab. You will not have time to complete most labs if you do not know the procedure first.

2. Do not remove stock containers of chemicals from their location. To take some of a chemical, pour the amount you need into a clean, dry, labelled container and take it to your work area. Do not take more than the amount in your instructions, and NEVER return chemicals to a stock container. Any leftover material should be thrown out at the end of the lab.

3. General Clean-up Rules:
   (a) All reusable equipment and glassware should be cleaned and returned to its proper place.
   (b) Solid wastes go into the waste basket; NEVER into a sink.
   (c) Liquid waste should be washed down the sink using lots of flush water.
   (d) Broken glassware and flammable liquids should be placed in special disposal containers.
   (e) Sinks should be rinsed and work area wiped clean at the end of each work period.
   (f) A good general rule is to leave your work area as you find it, but cleaner if possible.

4. Labels are important. Never take a chemical without checking the label on its container. Also, never put a chemical into a beaker or flask without first labelling it.

5. In general, never touch chemicals, put them in your mouth, or breathe their fumes unless so instructed. Wash immediately upon contacting any chemical. Wipe up any spills immediately and wash the area with water.

6. When heating material in a beaker, cover the beaker with a watch glass. When heating material in a test tube, point the test tube away from anyone in the area (including yourself). When heating glass, remember glass stays hot long after it looks cold, and glass burns are severe. Heated glass should be placed on wire gauze (NOT the bench surface) for cooling.

7. Report ANY accident, even a minor cut or burn, to your instructor IMMEDIATELY.

THE REACTION OF ZINC METAL WITH AQUEOUS LEAD(II) ACETATE. A SINGLE REPLACEMENT REACTION - LAB F1

Purpose: To determine stoichiometric relationships in the reaction between zinc metal and lead(II) acetate.

Apparatus and Materials:

1 - sample vial containing lead(II) acetate
1 - zinc metal strip
1 - centigram balance
1 - wash bottle containing distilled water
1 - 400 ml beaker
1 - 250 ml beaker
1 - watch glass
1 - piece of steel wool
1 - dropping bottle of glacial acetic acid.

Procedure:

Day One:(about 30 minutes)

1. Record the number of the centigram balance. All masses are to be determined to 0.01g.
2. Determine the mass of a clean dry 250 ml beaker.
3. Transfer the lead(II) acetate to the 250 ml beaker.
4. Determine the mass of the 250 ml beaker and the lead(II) acetate.
5. Rub the zinc strip with steel wool to clean its surface of oxides.
6. Determine the mass of the zinc strip.
7. Add about 200 ml of distilled water to the lead(II) acetate in the beaker.
8. Stir the mixture with a glass stirring rod until the lead(II) acetate crystals dissolve completely. Rinse off the stirring rod into the beaker as you remove it from the beaker using a wash bottle.
9. Add a few drops of glacial acetic acid to clear the solution.
10. Place the zinc strip into the lead(II) acetate solution and write visual observations for the first few minutes of the reaction.
11. Label the beaker with your initials, cover it with a watch glass, and put the beaker away until the next day.

Day Two: (about 20 minutes)

12. The reaction has now stopped. Write your visual observations about the contents of the beaker.

13. Remove the zinc strip from the solution, hold it over the beaker, and rinse any particles of lead (that may be sticking to the zinc strip) into the beaker using your wash bottle.
14. Set the zinc strip aside to dry completely. Determine the mass of the dry zinc strip.
15. Decant the solution into the 400 ml beaker as demonstrated by your teacher.
16. Wash the lead particles five times with approximately 10 ml portions of distilled water. Decant each washing into the 400 ml beaker.
17. Discard the contents of the 400 ml beaker.
18. Make sure the beaker is properly labelled and place it in a drying system until next day.

Day Three:

19. Determine the mass of the 250 ml beaker together with the lead.
20. Discard the lead, and clean up your station.

Results:

Day One:

Balance number __________
Mass of 250 ml beaker __________
Mass of 250 ml beaker and lead(II) acetate __________
Mass of zinc strip before reaction __________
Visual observations:

Day Two:

Visual observations:

mass of zinc strip after the reaction __________

6. The mass of Pb produced can be predicted by a stoichiometry calculation starting either from the amount of Pb $(CH_3COO)_2 \cdot 3H_2O$ that reacted or the amount of zinc that reacted.

   a. Predict the mass of Pb that should form in the reaction based on the amount of $Pb(CH_3COO)_2 \cdot 3H_2O$ that reacted.

   mass =

   b. Predict the mass of Pb that should form in your reaction based on the amount of Zn that was consumed by the reaction.

   mass =

   c. Determine the average of your results from (a) and (b) above; this is your predicted mass of lead.

   predicted mass=

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